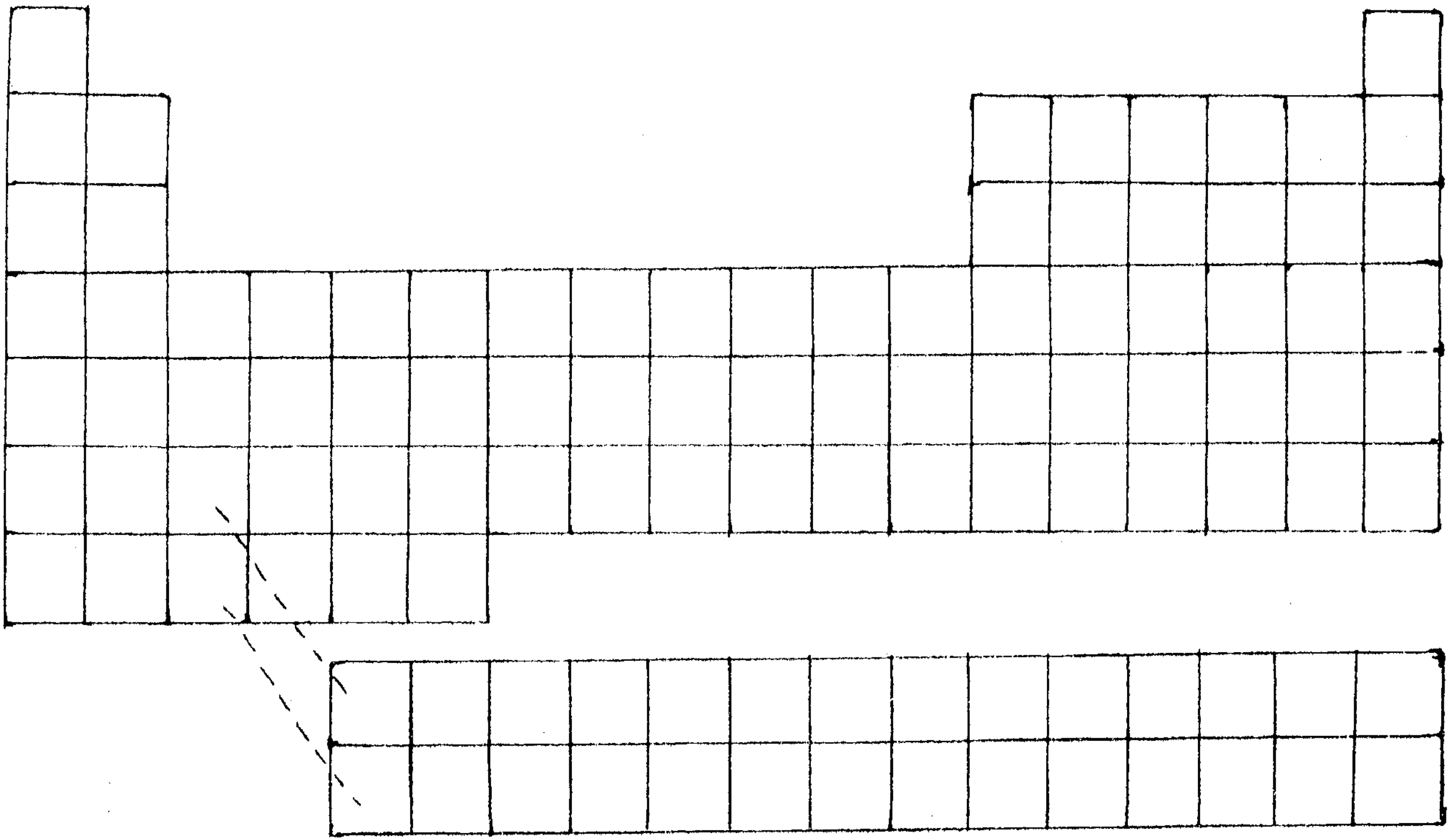
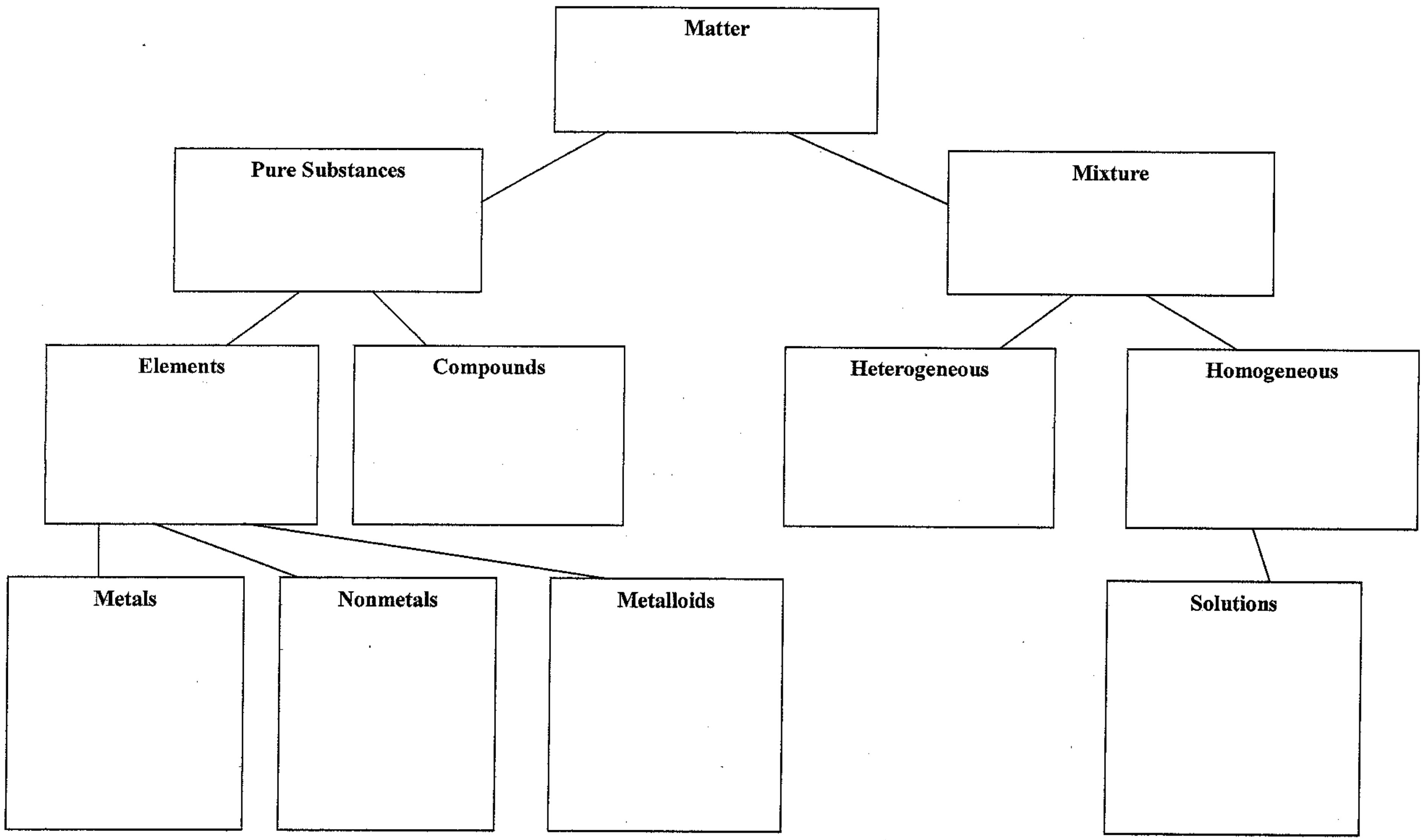
PERIODIC TABLE




Matter
Pure Substances
Mixture
Elements
Compounds
Heterogeneous
Homogeneous
Metals
Nonmetals
Metalloids
Solutions

Name ____________________

CHAPTER 4 WORK SHEET: MATTER

I. CLASSIFY THE FOLLOWING AS CHEMICAL OR PHYSICAL PROPERTIES

COLOR ____________
FLAMMABILITY ____________
POROSITY ____________
STABILITY ____________
EXPANSION ____________
RUSTING ____________
REACTIVITY ____________
ODOR ____________
REACTS WITH AIR ____________
DUCTILITY ____________
MELTING POINT ____________

II. ARE THE FOLLOWING PHYSICAL OR CHEMICAL CHANGES?

DIGESTION OF FOOD ____________
FADING OF DYE ON CLOTH ____________
GROWTH OF A PLANT ____________
FORMATION OF CLOUDS IN AIR ____________
HEALING OF A WOUND ____________
EXPLOSION OF GAS IN A CAR ENGINE ____________
MAKING ROCK CANDY ____________

III. CLASSIFY THE FOLLOWING AS HETEROGENEOUS, HOMOGENEOUS, COMPOUNDS OR ELEMENTS

AIR ____________
SALT ____________
ALCOHOL ____________
APPLE ____________
PAPER ____________
MILK ____________
PLUTONIUM ____________
SALT ____________

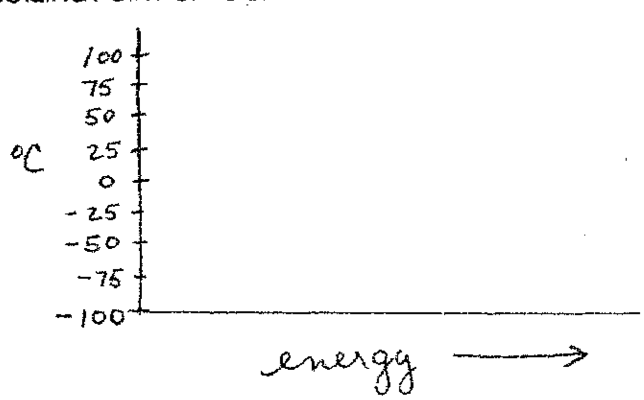
IV. HOW CAN YOU DECIDE IF THE FOLLOWING ARE HOMOGENEOUS OR HETEROGENEOUS?

A) A PIECE OF LUMBER ____________
B) A GLASS OF SODA ____________
C) CLOTH 50% WOOD AND 50% SYNTHETIC ____________
D) SHAVING CREAM ____________
E) MILK ____________

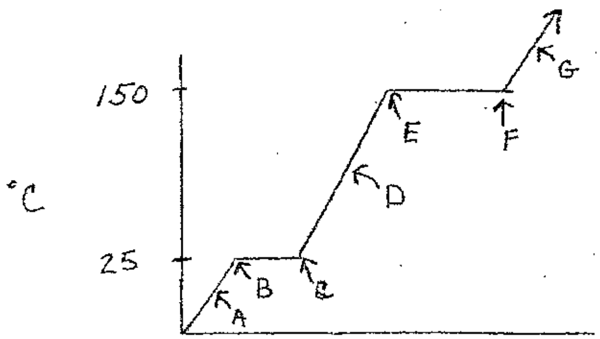
V. IDENTIFY THE FOLLOWING AS A COMPOUND, ELEMENT OR MIXTURE

A) WATER ____________
B) MILK ____________
C) SALT ____________
D) SILVER ____________
E) BRASS ____________
F) MAGNESIUM OXIDE ____________
G) ICE ____________
H) WOOD ____________

VI. DRAW A COMPLETE PHASE CHANGE GRAPH THAT REPRESENTS A MELTING POINT OF -50 C AND A BOILING POINT OF 75 C.



VII. USE THE FOLLOWING GRAPH TO ANSWER THE FOLLOWING QUESTIONS:



A) WHAT IS THE MELTING POINT?

B) WHAT IS THE BOILING POINT?

C) WHAT IS THE PHASE AT?

POINT A ______________
POINT B ______________
POINT C ______________
POINT D ______________
POINT E ______________
POINT F ______________
POINT G ______________

VII. CHANGE THE FOLLOWING TEMPERATURES TO EITHER °C OR K AS APPROPRIATE.

A) 200 C = _______K

B) -120 C = _______K

C) 0 C = _______K

D) 500 C = _______K

E) -273 C = _______K

F) 50 K = _________ C

G) 100 K = ________C

H) 2000 K=________C

I) 1K = ________C

J) 273 K = ________C

Name ____________ Class ____________ Date ____________

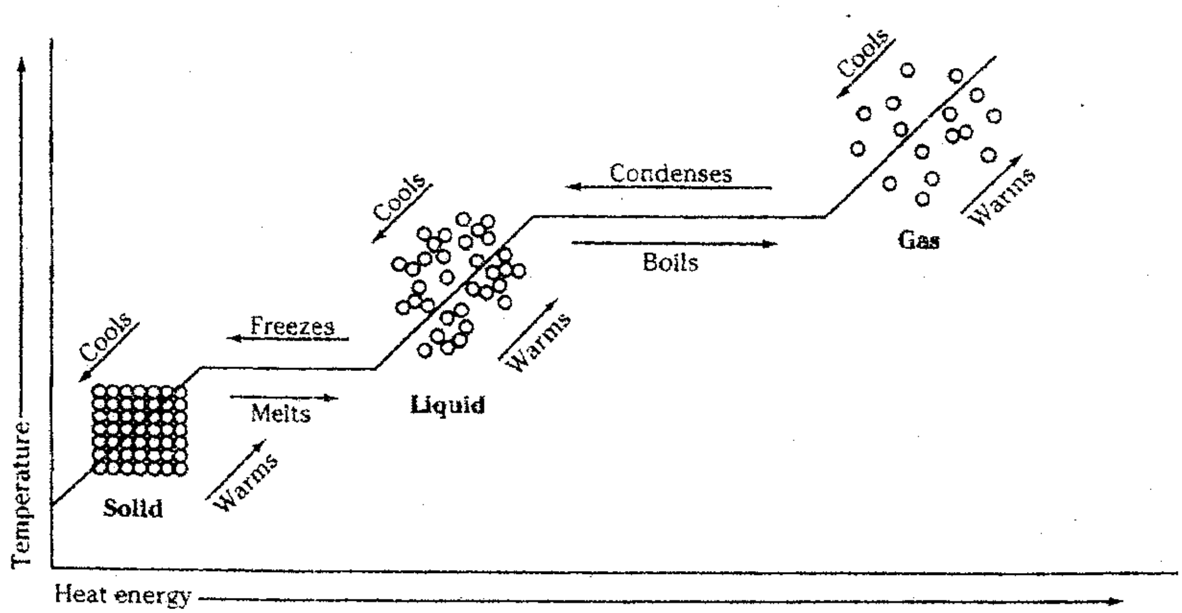
CHAPTER 3

Introduction to Chemistry
Section 3–1

SKILL ACTIVITY
Interpreting graphs

Observing Phase Changes

The accompanying graph shows the relationship between temperature and heat energy during the phase changes of water. Study the graph and answer the following questions.



1. Does the temperature increase during melting? ____________

____

2. Is energy required for each phase change? ____________
3. Can both liquid water and steam exist at 100°C? ____________
4. What must be changed—temperature or heat energy—during condensation? ____________
5. How would you describe the change in the arrangement of particles as heat energy and temperature increase? ____________

____

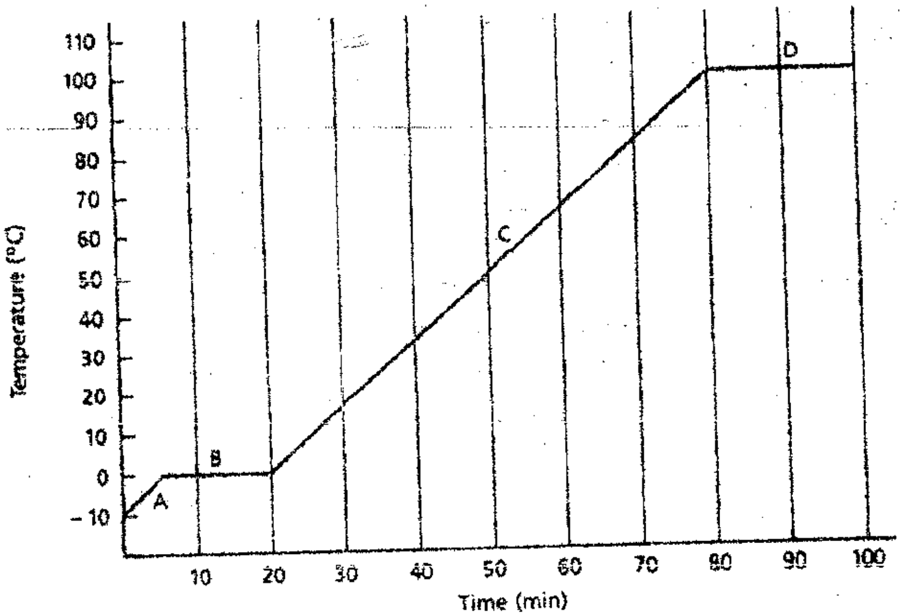
6. What rule can you state about the relationship between phase changes and temperature? Between changes and heat energy? ____________

____

Name ______________________ Date __________ Class __________

A Heating Curve

The heating curve shown in the figure is a plot of temperature *vs.* time. It represents the heating of what is initially ice at −10°C at a constant rate of heat transfer.



Answer the following questions.

1. a. What phase or phases are present during segment A? **1. a.** ______________________

b. What is happening to the energy being absorbed from the heat source? (Answer in terms of potential and/or kinetic energy.) **b.** ______________________

c. What phase change, if any, is taking place? **c.** ______________________

2. a. What phase or phases are present during Segment B? **2. a.** ______________________

b. What is happening to the energy being absorbed? **b.** ______________________

c. What phase change, if any, is taking place? **c.** ______________________

d. What is the significance of the temperature 0°C? **d.** ______________________

A Heating Curve (continued)

3. a. What phase or phases are present during segment C? 3. a. ____________________

b. What is happening to the energy being absorbed? b. ____________________

c. What phase change, if any, is taking place? c. ____________________

4. a. What phase or phases are present during segment D? 4. a. ____________________

b. What is happening to the energy being absorbed? b. ____________________

c. What phase change, if any, is taking place? c. ____________________

d. What is the significance of the temperature 100°C? d. ____________________

5. What would you expect to happen if the heating were continued? 5. ____________________

____

____

____

Chemistry Lab #2 – Melting Point/Freezing Point of a Pure Substance

Procedure:

1.) Obtain a test tube of stearic acid, a computer, and a computer pack. Remove the temperature probe, the USB cable, and the universal laboratory interface unit from the computer pack. Connect the USB cable to the interface unit and the computer. Connect the temperature probe to channel 2 on the interface unit. Turn on the computer, log on, and open the program for the melting point lab.

2.) If the temperature probe is already in the test tube, you will be determining the melting point of the stearic acid. First you need to heat a beaker of water to about 70°C with a Bunsen burner. **Make sure the temperature probe wire stays away from the flame!** When the water is heated, place the test tube in the beaker and secure the test tube with a buret clamp. Set the collection time for 1200 seconds. Begin collecting data and continue collecting for 1200 seconds or until the stearic acid is completely melted. Note when the stearic acid begins to melt and when it has melted completely. Title the graph "Solid to Liquid" and remember to label the x-axis and y-axis. Print out 2 copies of the graph **only**, not all the data that goes with it.

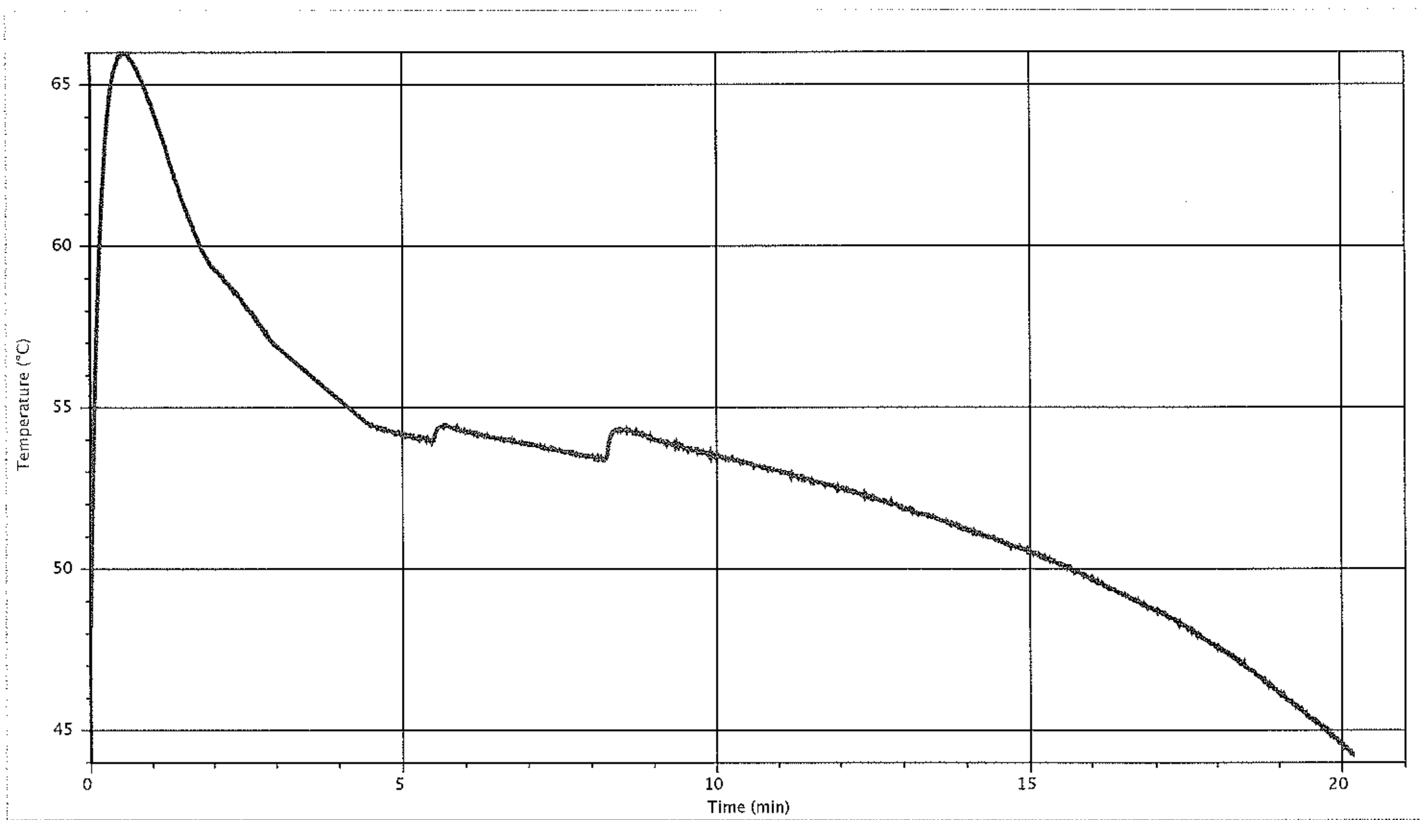
OR

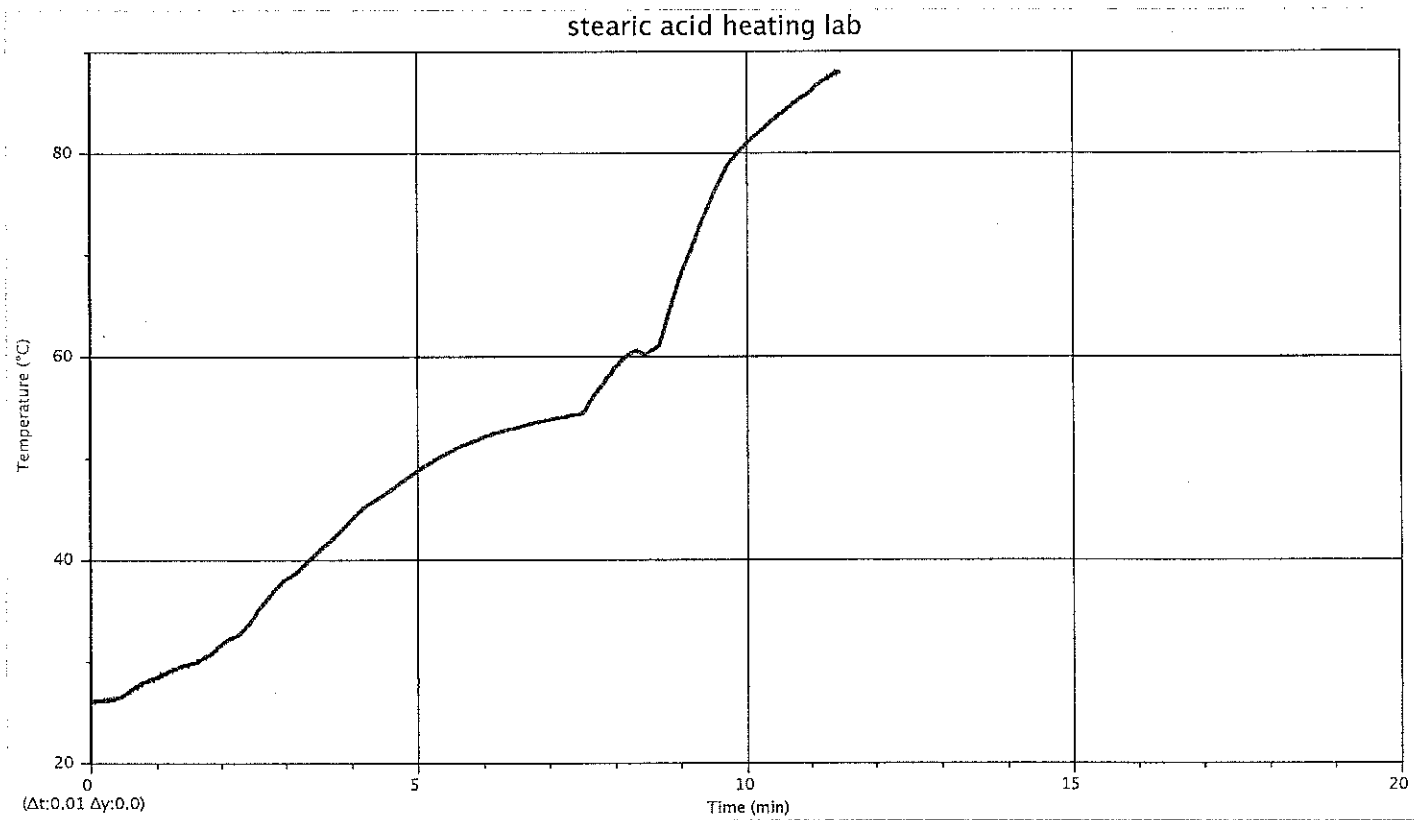
2.) To determine the freezing point of the stearic acid, you will need to melt the stearic acid. To melt the stearic acid, place the test tube in a beaker of water. Clamp the test tube with a buret clamp. **Be sure to remove the cork before heating.** (beaker should be about 2/3 full). Heat the water to ~70°C. You do not want the water to boil. When the stearic acid is completely melted, place the temperature probe into the test tube. Remove the test tube from the water and turn off the Bunsen burner. Set the collection time for 1200 seconds. Start collecting data when the temperature probe in the stearic acid has reached about 70°C and is coming down. Continue collecting data for 1200 seconds or until the substance is completely solidified. Note when the stearic acid begins to solidify and when it has solidified completely. Title the graph "Liquid to Solid" and remember to label the x-axis and y-axis. Print out 2 copies of the graph **only**, not all the data that goes with it.
To remove the probe from the stearic acid after it has solidified, heat the test tube in the beaker of water. **Keep the temperature probe wire away from the flame**. When the stearic acid has melted, remove the probe and wipe clean.

3.) Place the test tube of stearic acid back in the location specified by your teacher. Make sure the temperature probe is clean and put the computer and the accessories back in the computer cart.

Discussion Questions

1.) Based on your observations, what is the melting point of the stearic acid? What is the freezing point of the stearic acid?
2.) Describe the shape of the two graphs. What do you think is happening during the time where the flat spot occurs?
3.) Is energy still being transferred to the stearic acid? If so, what is that energy doing?
4.) What would a complete temperature vs time graph look like if it were done perfectly?


65
60
55
50
45
Temperature (°C)
0
5
10
15
20
Time (min)



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NAME________________ **Specific Heat Problems**

1. How much heat would be needed to raise the temperature of 100 mL of water from 20^0C to 80^0C ?

2. If 20,900 J of heat energy is added to a 70 g. piece of steel at 20^0C, what is the final temperature? Cp = .418 J/g°C.

3. If 4,180J of heat energy raises the temperature of a piece of aluminum (Sp.H.) 0.8778 J/g^0C from 0^0C to 45^0C, what is the mass of the aluminum?

4. If 41,800 J of heat energy is added to 500 mL of water at 20^0C, what is the final temperature?

5. If 20,900 J of heat energy raises the temperature of 200 g of an unknown substance from 20^0C to 90 ^{0}C, what is the specific heat capacity of the metal?

6. How much heat would be required to heat 390 kg of nickel from 25 ^{0}C to 300^0C? (Cp of Ni = 0.443 J/g^0C)

7. How much heat would be required to heat 54.4 g of tin from 24 ^{0}C to 100 ^{0}C? (Cp Sn= .220J/g^0C)

8. If a piece of cobalt with a mass of 25.2 g and a temperature of 100 ^{0}C is dropped in 15 cm^3 of water at 22 ^{0}C, what will be the final temp o the system? (Cp Co= 0.446 J/g^0C)

9. If a piece of unknown metal with a mass of 32.2g is heated to 100^0C and dropped in 60 mL of H_20 at 25^0C, what is the specific heat capacity of the metal if the final temperature of the system is 35.2 ^{0}C?

10. Calculate the energy needed to raise 30g of ice from −10°C to 125°C

HEAT WORKSHEET#2

1. 2700 cal of heat energy is added to 200 ml of acetone that has a temperature of 20°C. What is the final temperature? (Acetone has a density of 0.79 g/ml and a specific heat capacity of 0.52 cal/g°C).

2. How much heat is required to raise the temperature of one liter of water from 20°C to 25°C?

3. Steel has a specific heat capacity of 0.107 cal/g°C while aluminum's is 0.21 cal/g°C. If you had two pots of equal mass, 1.5 kg, which pot would take more heat to heat up AND how much more heat would it take to heat the one pot from 20°C to 220°C?

4. The melting point of brass is 900°C. How much heat would be needed to get 800 g of brass from 20°C to the melting point? (Specific heat capacity of brass is 0.092 cal/g°C)

5. 2600 cal of heat energy is added to a piece of copper at 27°C. The temperature rises to 53°C. What is the mass of the copper? (Specific heat of copper is 0.091 cal/g°C)

HEAT PROBLEMS CONT'D

1. How much heat would be needed to take 500 g of steel from -20°C to 75°C? (Specific heat capacity of steel is .11 cal/g°C)

2. If you add 50000 cal to 80 g of water at -75°C, what will the final temperature and state be of water?

3. 500 cal of heat causes a piece of aluminum to go from 20°C to 55°C. What is the mass of the Al? (Specific heat capacity is .21 cal/g°C)

4. 250 g of an unknown metal at 100°C is placed in 150 ml of water that is at 20°C. When equilibrium is reached, the temperature is 23°C. What is the specific heat capacity of the metal?

5. 2000 cal of heat is added to 90 g of brass at 20°C. What is the final temperature of the brass? (Specific heat capacity is .092 cal/g°C)

6. How much heat would be needed to take 100 g of ice at -58°C to liquid water at 46°C?

100

Chemistry Practice problems Heat Equilibrium

1. A 110 g piece of steel at 20°C is placed in 200 ml of H_2O at 100°C. What is the final temperature of the metal and water? (SHC steel use .107 cal/g°C)
2. A student heats a 150g piece of aluminum to 1350°C and then places it in cool 20°C water to cool it off. If the student uses 100 ml of water, will it still be too hot to touch? (SHC aluminum use .84 J/g°C)
3. How much 20°C water would be needed to cool a 1 Kg horseshoe from 1750°C to 50°C so that it can be picked up without burning the user? (SHC of metal shoe is .11cal/g°C)
4. Describe how you would measure the specific heat capacity of an unknown metal.
5. A 250 g piece of unknown metal at 100°C is placed in 150 ml of water at 25°C. The equilibrium temperature is 29.3°C. What is the specific heat capacity of the metal?

100

Chemistry Homework   Heat Problems

1. How much heat would be needed to raise the temperature of one liter of water from its freezing point to its boiling point?
2. A 50 g piece of steel has 1500 cal of heat energy added to it. If the specific heat capacity of steel is .448 J/g°C and the original temperature is 20 °C, what is the final temperature? (1 cal = 4.184 J)
3. 2000 Joules is added to a piece of steel and the temperature goes from 20 °C to 500 °C. What is the mass of the steel? Same specific heat capacity as the last problem.
4. Define specific heat capacity and calorie.
5. 5000 cal of heat energy is added to a 300g piece of aluminum, (specific heat capacity .21cal/g°C), that has a temperature of 20 °C. What is the final temperature of the aluminum?
6. How much heat would be needed to change 50g of ice at -40°C to liquid water at 20°C? (specific heat capacity of ice is .5cal/g°C)(Heat of fusion of water is 80cal/g)
7. How much heat would be needed to melt 100g of ice at its melting point?
8. Draw a melting point curve and label the different sections.
9. Describe what occurs during a phase change.
10. What is the melting point of water? What is the freezing point?

Chemistry Lab #3 Specific Heat Capacity of Metals

Procedure:

1) You will need to set up a hot water bath that will be used to heat the metals to 100 C. Start the water bath with your bunsen burner as soon as you begin lab. Use your 400 ml beaker and don't forget the wire pad.
2) Obtain one of the metals that is to be used. (only take one at a time) Measure its mass on the triple beam balance and record the value in your data table.
3) After you have measured the mass, hang the block with a piece of wire from your buret clamp and clamp it to the ring stand so that the metal is suspended in the hot water bath. The metal should not rest on the bottom of the beaker, but it is alright if it touches the sides. <u>Be very careful not to knock the beaker of hot water over</u>. Your teacher will demonstrate this in class.
4) After the water has begun to boil, measure its temperature with your thermometer. This will be the only time you will need to measure the temperature of the water bath. We will assume the temperature will remain constant for the rest of the metals. Record this temperature in your data table under the original temperature of the metal. (Remember, the metal will be in the water for 5 minutes and will have the same temperature as the water)
5) Obtain a styrofoam cup and place 100 ml of tap water into it using a 100 ml graduated cylinder. (This water will need to be changed for each metal)
6) After the metal has been in the boiling water for about 5 minutes, you are going to transfer it to the water in the styrofoam cup. This procedure will need to be done as a team. One partner will read the temperature of the water in the styrofoam cup JUST before the other partner transfers the metal from the boiling water. This will be recorded in your data under original temperature of the water. Transfer the metal as quickly and as safely as possible. Watch the temperature of the water as it rises. Record the highest temperature reached in your data under final temperature of the metal and water.
7) One hint that will save you some time will be to get the next metal block and weigh it while you are waiting for the last metal to heat in the boiling water.
8) Repeat the steps above with all four metals. Remember to change the water in the styrofoam cup each time.

Questions:

1) Complete a sample calculations section for one of the metals. Your teacher will show you how this is done.
2) Give a definition for the calorie and specific heat capacity.

3) What does the specific heat capacity tell you about the substance.
4) Complete the following statement: The lower the specific heat capacity is, the ?????
5) If 50 g of steel that had a temperature of 100 C was placed in 50 g of water with a temperature of 0 C, why dosen't the temperature end up at 50 C (half way between 100 and 0 C).

Data Table.

metal	mass(g)	metal T_i (°C)	metal + H_2O T_f (°C)	H_2O T_i (°C)	metal C_p ($\frac{J}{g°C}$)	Actual C_p ($\frac{J}{g°C}$)	% err
aluminum						.900	
brass						.385	
steel						.418	
lead						.138	
H_2O	100ml = 100g	X	X	X	X	4.18	X

C_p of metals

Group	metal	mass (g)	metal T_i (°C)	metal + H_2O T_f (°C)	H_2O T_i (°C)	exp C_p (J/g°C)	theor C_p (J/g°C)	% error
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	
	Al						.900	
	Brass						.385	
	Steel						.418	
	Pb						.138	

I Title

II Purpose -use complete sentences

III Data table- Get it Signed!

Data table: Specific Heat Capacity of Metals

metal	mass (g)	initial temp of metal (°C)	final temp metal + H_2O	initial temp of H_2O (°C)	C_p exp (J/g-°C)	C_p actual (J/g-°C)	% error
aluminum						.900	
brass						.385	
steel						.418	
lead						.138	
H_2O	100ml= 100g					4.18	

IV Calculations. Show All work!

A) Show work for C_p exp for all 4 substances

B) Show work for % error for all 4 substances.

A. heat lost (metal) = heat gained (H_2O)

$$m(\Delta T)C_p = m(\Delta T)C_p$$

B. $\% \text{ error} = \left| \dfrac{\text{Exp} - \text{Theor}}{\text{Theor}} \right| \times 100$

V Analysis. Questions 1→5

1) See calculations section
2) calorie –
   C_p –
3)
4)
5) Do the calculation. $m(\Delta T)C_p = m(\Delta T)C_p$ with the values given

VI Conclusion - use complete sentences

A. Summarize the procedure
B. What did you learn?
C. Explain how you know if the purpose was accomplished.

Lab Write-up for Specific Heat of Metals

1. **Title**
2. **Purpose**
3. **Method**
4. **Data Table:** 9 columns

Metal	Mass (g)	Volume H_2O (mL)	Metal T_i (°C)	Metal+H_2O T_f (°C)	H_2O T_i (°C)	Exp. C_p (J/g°C)	Theor. C_p (J/g°C)	% Error
Al							0.900	
Brass							0.385	
Steel							0.418	
Pb							0.138	

Note: Cp water = 4.184J/g°C

5. **Introduction to Calculations:**
   - Two calculations: experimental Cp and % error
   - This is a brief description of what they are going to show. Explain the equation in words. This is not a list of steps that you will follow. For example, in this experiment, the specific heat capacity of four metals was determined by combining the hot metal at a known temperature with water at a known temperature, and then measuring the final temperature once equilibrium has been reached. In equation form it would be …..where Cp is the specific heat capacity of the metal, m is mass, q is the heat gained/lost, and ΔT is the change in temperature.

6. **Sample Calculations:**
   1. To determine specific heat capacity……This is a very detailed step-by-step description of the equation.
   2. Show one example for each type of calculation (Cp and % error).

      Specific Heat Capacity:

      Q (heat lost by metal) = Q(heat gained by water)

      $m\Delta TCp = m\Delta TCp$

Data:

Steel	Water
m = 249.1g	m = 100g (100mL)
Ti = 95°C	Ti = 23°C
Cp = ?	Cp = 4.184 J/g°C

Tf = 37°C

(249.1g)(58°C)Cp = (100g)(14°C)(4.184 J/g°C)

Cp = 5857.6/14447.8 = 0.405 J/g°C

Percent error: (actual – theoretical)/theoretical x 100

(0.405 J/g°C-0.418 J/g°C)/0.418 J/g°C x100 = -3.11%

7. **Discussion:** This is the section where you will convince me that you understand all the concepts in the lab. You want to define concepts here, explain things, and talk about your answers.
    - Answer discussion questions here.
        1. Sample calculations
        2. Define calorie and specific heat capacity – explain in words, what is the difference between c and C?
        3. What does Cp tell you about the substance? Use examples, don't limit these to the experiment, think about examples used on class – aluminum foil, the ocean, sand, boiling water, metal pot. Can talk about the relationship between density and Cp, relate labs to one another
        4. Complete sentence: The lower the Cp is, the ?????? You should include <u>several</u> statements here; explain.
        5. If 50g of steel that had a temperature of 100 °C was placed in 50g of water with a temperature of 0°C, why doesn't the temperature end up at 50°C (half way between 100°C and 0°C)? Explain process, solve the problem (real final temp is 9.08°C)
    - Address three possible errors – be sure to state specifically how the error affects the experimental Cp value.
    - Additional Information:
        1. Information about the metals
        2. Information about mixtures and pure substances
        3. Heterogeneous vs. homogeneous
        4. Characteristics/properties of metals
        5. General statements about Cp – What substances have high/low Cp? What does that mean? Give examples.
        6. Relationship between J and cal.
        7. Show a graphical relationship:

            heat lost by the metal =heat gained by the water

Experiment 14: Energy Needed to Melt Ice

Introduction:

Pure substances have characteristic melting and freezing behavior. Pure water changes from solid (ice) at 0°C to a liquid at 0°C as energy is added to it. In this experiment, you will determine the energy required to melt one gram of ice by letting an excess of it interact with warm water in a Styrofoam cup (ice calorimeter). The ice will cool the water to about 0°C. The energy given up by the water as it cools is the energy used to melt the ice. Recall that one-calorie is required to change the temperature of one gram of liquid water one-Celsius degree.

Standards:

3.2.10C Apply elements of scientific inquiry to solve problems
3.4.10B Apply appropriate instrument and apparatus to examine a variety of objects and processes
3.4.10b Analyze energy sources under which heat is transferred

Materials:

Ice and water
400 mL beaker
100 mL graduated cylinder
Bunsen burner
Styrofoam cup
Ringstand/ring/wire gauze
Thermometer
Tongs

Procedure:

1. Warm approximately 250 mL of water to about 55°C.
2. Measure 100 mL of luke warm water into a Styrofoam cup. Record the temperature of the warm water to the nearest whole number.
3. Obtain several ice cubes. Shake excess water from them. Place the ice in the warm water and stir the mixture until the temperature is about 0°C. Add more ice if needed to cool the water. Record the lowest temperature reached.
4. After the water and ice mixture has cooled to about 0°C, remove the unmelted ice (using tongs). Be sure to drain off as much water as possible into the cup when you remove the ice.
5. Measure the volume of water remaining in the calorimeter to the nearest milliliter.

Calculations:

1. Calculate heat of fusion of ice using trial data #1. (cal & J)
2. Calculate heat of fusion of ice using trial data #2. (cal & J)
3. Calculate heat of fusion of ice using trial data #3. (cal & J)
4. Calculate average trials 1, 2, 3. (cal & J)
5. Calculate percent error on the average. (cal & J)

Questions:

1. In what way does calorimetry make use of the Law of Conservation of Energy?
2. What is the difference between heat and temperature?

Data – Heat of Fusion

1. Trial # _______
2. Volume of heated water: __________
3. Temperature of heated water: ___________
4. Temperature of final mixture: ___________
5. Change in temperature of water: __________
6. Volume of heated water + ice (water): _________
7. Mass of ice (water): ___________

1. Trial # _______
2. Volume of heated water: __________
3. Temperature of heated water: ___________
4. Temperature of final mixture: ___________
5. Change in temperature of water: __________
6. Volume of heated water + ice (water): _________
7. Mass of ice (water): ___________

1. Trial # _______
2. Volume of heated water: __________
3. Temperature of heated water: ___________
4. Temperature of final mixture: ___________
5. Change in temperature of water: __________
6. Volume of heated water + ice (water): _________
7. Mass of ice (water): ___________

H_{fus} Lab

theor (79.71 cal/g or 333.51 J/g)

Group	trial	V heated H_2O	T_i heated H_2O	T_f mixture	ΔT H_2O	V heated H_2O + ice mix	V ice	Exp H_{fus}	% error

Lab Write-up for Heat of Fusion

1. **Title**
2. **Purpose**
3. **Method**
4. **Data Table:** 9 columns

Trial	Volume H_2O (mL)	Temp of heated H_2O (°C)	Final Temp (°C)	Temp Change of H_2O (°C)	Volume of H_2O + ice (mL)	Mass of ice (water) (g)	Exp. H_f (cal/g)	Exp. H_f (J/g)
1								
2								
3								

Average Experimental H_f –________ cal/g
Average Experimental H_f –________ J/g
Theoretical H_f – 79.71 cal/g
Theoretical H_f – 333.51 J/g
Percent Error (cal/g) -______________
Percent Error (J/g) -_______________

5. **Introduction to Calculations:**
   - Four calculations: experimental H_f in cal/g and J/g, average H_f, and % error
   - This is a brief description of what they are going to show. Explain the equation in words. This is not a list of steps that you will follow.
     1. **Heat of Fusion (cal/g)**: In this experiment, the heat of fusion for H_2O was determined. Knowing that the heat lost by the warm water is equal to the heat gained by the ice water plus the heat needed to melt the ice, heat of fusion for H_2O can be calculated. In equation form it would be ….where Cp is the specific heat capacity of water, m is mass, ΔT is the change in temperature, and H_f is the heat of fusion for H_2O.
     2. **Heat of Fusion (J/g)**
     3. **Average Heat of Fusion**
     4. **% Error**

6. **Sample Calculations:**
   1. To determine heat of fusion for H_2O ……This is a very detailed step-by-step description of the equation.
   2. Show one example for each type of calculation (H_f (cal/J), H_f (J/g), average H_f and % error).

   <u>H_f for H_2O (cal/g):</u> (Draw graph on board)

   Q (heat lost by warm H_2O) = Q (heat gained ice H_2O) + Q (H_f H_2O)

   $$m\Delta TCp = m\Delta TCp + mH_f$$

Data: Trial 1

Warm Water	Ice Water
$m = 99$ g	$m = 52$g
$Ti = 53°C$	$Ti = 0°C$
$Cp = 1$ cal/ g°C	$Cp = 1$ cal/g°C

$Tf = 3°C$

$$(99g)(50°C)(1\ cal/g°C) = (52g)(3°C)(1\ cal/g°C) + 52g(H_f)$$
$$4950\ cal = 156\ cal + 52g(H_f)$$
$$4794\ cal = 52g(H_f)$$
$$4794\ cal/52g = H_f$$
$$92.1923\ cal/g = H_f$$

Significant Figures: 92 cal/g

H_f calories to Joules: 92 cal/g x 4.184 J/cal = 384.028 J/g (380 J/g)

Average: Add three experimental values and divide by three

Percent error: (experimental – theoretical)/theoretical x 100

7. **Discussion:** This is the section where you will convince me that you understand all the concepts in the lab. You want to define concepts here, explain things, and talk about your answers.
   - Answer discussion questions here.
     1. In what ways does calorimetry make use of the Law of Conservation of Energy? What is the law? What does the law mean? How does the law relate to this lab?
        Use the equation: Q (heat lost) = Q (heat gained) + Q (H_f)
        Draw the graph – LABEL properly (title, axes, units)
     2. What is the difference between heat and temperature?
        Provide definitions of each, tell me the difference between the two
   - Address three possible errors – be sure to state specifically how the error affects the experimental H_f value.
   - Additional Information:
     1. Characteristics/differences between ice and water (H_2O liquid and H_2O solid) Can talk about all solids and liquids – compare and contrast
     2. Relate to other labs – density, Cp, MP/FP (physical properties and phase changes)
     3. Relationship between J and cal, Calories and calories
     4. Where did the name "Joule" come from?
     5. Kinetic energy vs. potential energy and how they relate to this lab

Advanced Chemistry

Series I Number 9

Heat of Fusion of Water

Pure substances have characteristic melting and freezing behavior. Pure water changes from a solid at 0 degrees celsius to a liquid as energy is added to it. In this experiment you will determine the energy required to melt one gram of ice by letting an excess of it interact with warm water in a styrofoam cup. The ice will cool the water to about zero degrees celsius. The energy given up by the water as it cools is the energy used to melt the ice. Recall that one calorie is required to change the temperature of one gram of liquid water one degree celsius.

Procedure:

1) Warm about 125 ml of water to about 50°C.
2) Measure 100 ml of this warm water within 1 ml into a styrofoam cup. Record the temperature of the warm water to the nearest 0.1°C. (Read the temperature just before you do the next step).
3) Obtain some ice cubes. Shake excess water from them. Place the ice in the warm water and stir the mixture until the temperature is below 5°C. It is very important that this process only take 40 or 50 seconds. If it is taking longer, repeat the experiment using a larger amount of ice. Do not attempt to get to 0°C. It will take too long and too much energy is lost to the surroundings.
4) Record the lowest temperature reached. Quickly remove any ice that remains using your fingers. Be sure to drain back as much water as possible into the cup when you remove the ice.
5) Measure the volume of water remaining in the cup to the nearest milliliter.

Calculations:

1) Determine the change in temperature of the warm water in the cup.
2) Using, $Q=m\Delta tC_p$, determine the amount of heat lost by the warm water as it underwent the temperature change in #1. (Density of water 1 g/cm^3, specific heat capacity 1 cal/g°C).
3) Since you know the initial volume and final volume, (after the ice melted) you can determine the mass of ice that melted. (Remember, water is 1 g/cm^3).
4) Not all of the energy lost by the warm water was used to melt the ice. After it melted (at 0° C) it then went up in temperature to the temperature that the warm water fell to. Use, $Q=m\Delta tC_p$, to determine the amount of energy gained by the ice (after it melted) as it went from 0°C up to the temperature the warm water fell to. (Make sure you use the mass of ice, not the mass of the warm water).
5) Determine the amount of energy that melted the ice. To do this subtract the amount

of energy found in #4 from the total energy lost by the warm H_2O, #2.

#2 (energy lost by warm water) - #4 (energy gained by small temp change after ice melted) = ______energy that was needed the melt the ice.

6) Determine the heat of fusion of water. Divide the energy found in #5 by the mass of ice melted. This represents the number of calories needed to melt 1 g of ice, or the heat of fusion.

7) The actual value is 79.71 cal/g. Do a % error.

Questions:

1) Define melting point and heat of fusion.
2) Draw a complete melting point, boiling point curve for a general substance. (temperature vs time).
3) Using the concepts of the kinetic molecular theory, describe the difference between solid, liquids, and gases in terms of the molecules.
4) Do all solids have melting points? Explain
5) Water has a very high heat of fusion and melting point considering the small molecular weight. Why is this so?
6) Water is less dense as a solid then a liquid. In terms of the molecule, why is this true.
7) List possible experimental errors and tell how each one effects your answer. (at least 3).

Chemistry Lab #4 - Heat of Combustion for Candle Wax

Procedure:

Part I

1) Obtain a candle and an index card. Attach the candle to the index card by melting some wax onto the card and placing the candle into the hot wax. Weigh the card and candle together on the centigram balance to the nearest 0.01g.
2) Weigh an empty can on the balance. After you have weighed the can, add tap water so that it is about two-thirds full. Do not measure the mass or volume of the water at this time.
3) Set up the apparatus as shown by your teacher so that the flame of the candle when lit (do not light it yet) will almost but not quite touch the bottom of the can.
4) Cool the water with ice, if necessary, fo that its temperature is about 10-15 degrees below room temperature. Add the ice directly to the water. Remove any remaining ice when the desired temperature has been reached.
5) Read and record the temperature of the water to the nearest 0.2 degrees. Light the candle and heat the water, stirring it gently, until it reaches a temperature about as much above room temperature as it was below at the start. Carefully blow out the candle flame. Continue to stir the water, while watching the thermometer reading, until the highest temperature is reached. Record the highest temperature to the nearest 0.2 degrees.
6) Read and record the mass of the candle and index card. Read and record the mass of the can and water.

Part II

1) Obtain a test tube that is filled with wax and a test tube of the same size that is empty. Read and record the mass of both test tubes.
2) Place 200ml of water into your 250 ml beaker using a 100ml graduated cylinder.
3) Heat a water bath with your bunsen burner and place the test tube that contains the wax into it so that the wax melts.
4) Using a test tube holder, remove the test tube containing the melted wax from the hot water. Hold the test tube in the air until the first sign of cloudiness (solidification) is evident. While the wax is cooling, measure and record the temperature of the water in the beaker. Quickly place the test tube into the water in the beaker, and record the highest temperature reached.

Questions

1) Calculate the heat of combustion and heat of fusion of candle wax. (Your teacher will demonstrate this to you)
2) What is the difference between burning and melting?
3) How does the heat of combustion compare with the heat of fusion.
4) When the candle burns, what happens to the wax? Could you mold the melted wax back into a candle of the same size?

I Title : Heat of Combustion for Candle Wax

II Purpose - use complete sentences

III Data - Get it Signed!

Part I

	Before	After
card + candle	g	g
empty can	g	
H_2O	°C	°C
can + H_2O		g

Part II

test tube + wax	______ g
test tube	______ g
H_2O	200 ml = 200 g
H_2O in beaker at condensation	______ °C
H_2O in beaker at fusion	______ °C

IV Calculations (Look at p 283-284, Section 11:15).

Show work for calculating H_c (heat of combustion) and H_f (heat of fusion)

$$\overbrace{m \quad H_c}^{wax} = \overbrace{m \quad \Delta T \quad C_p}^{H_2O}$$

mass: before candle − after candle = wax ; heat of combustion = can + H_2O − can = H_2O g ; After °C − Before °C ; 4.18 J/g°C

$$\overbrace{m \quad H_f}^{wax} = \overbrace{m \quad \Delta T \quad C_p}^{H_2O}$$

mass: T + wax − TT = wax ; heat of fusion = 200 g ; fusion °C − condensation °C = ΔT ; 4.18 J/g°C

V Analysis Questions

1) See calculations section
2) define burning -
   melting -
3) define H_c -
   H_f -
4) think questions!

VI Conclusion - use complete sentences!

A - Summarize the procedure
B - Tell me what you learned
C - Did you accomplish the purpose? Explain!

Lab: Candle Wax

Group	Pt 1: H_c			Pt 2: H_f	
		Before	After		
	card + candle	g	g	test tube + wax	g
	empty can	g	x g	test tube	g
	H_2O	°C	°C	H_2O	200ml=200g
	can + H_2O	x g	g	H_2O in beaker at condensation	°C
				H_2O in beaker at fusion	°C
	card + candle	g	g	test tube + wax	g
	empty can	g	x g	test tube	g
	H_2O	°C	°C	H_2O	200ml=200g
	can + H_2O	x g	g	H_2O in beaker at condensation	°C
				H_2O in beaker at fusion	°C
	card + candle	g	g	test tube + wax	g
	empty can	g	x g	test tube	g
	H_2O	°C	°C	H_2O	200ml=200g
	can + H_2O	x g	g	H_2O in beaker at condensation	°C
				H_2O in beaker at fusion	°C
	card + candle	g	g	test tube + wax	g
	empty can	g	x g	test tube	g
	H_2O	°C	°C	H_2O	200ml=200g
	can + H_2O	x g	g	H_2O in beaker at condensation	°C
				H_2O in beaker at fusion	°C

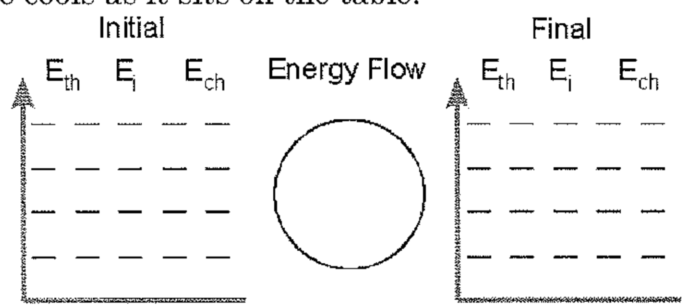
Name________________________________

Date__________ Pd____

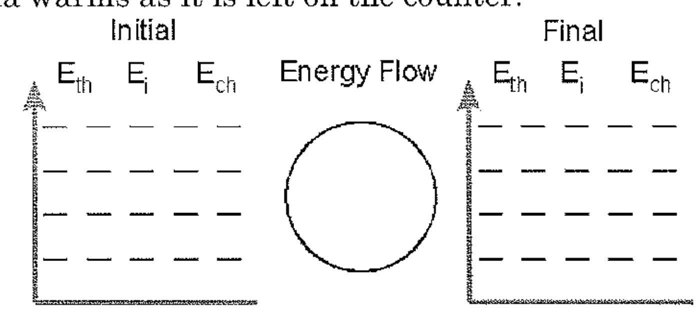
Unit 3 - Worksheet 1

For each of the situations described below, use an energy bar chart to represent the ways that energy is stored in the system and flows into or out of the system. Below each diagram describe how the arrangement and motion of the molecules change from the initial to the final state.

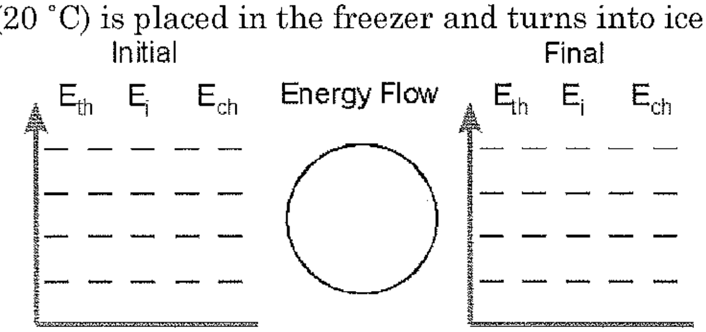
1. A cup of hot coffee cools as it sits on the table.



2. A can of cold soda warms as it is left on the counter.

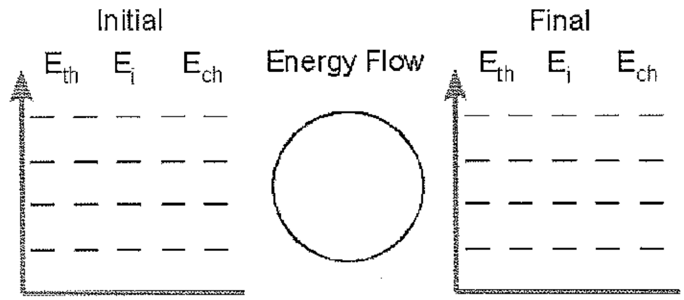


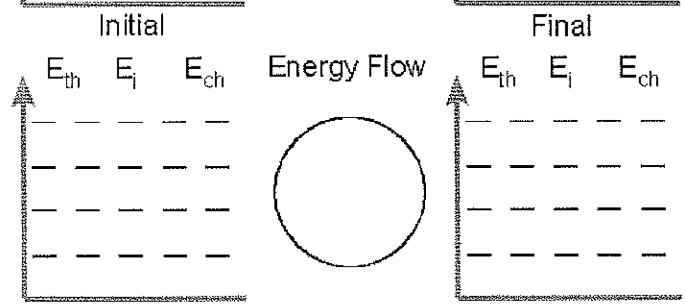
3. A tray of water (20 ˚C) is placed in the freezer and turns into ice cubes (- 8 ˚C)



4. Where does the energy that leaves the system in #3 go? How does this energy transfer affect the room temperature in the kitchen? Do you have any experience that supports your answer?

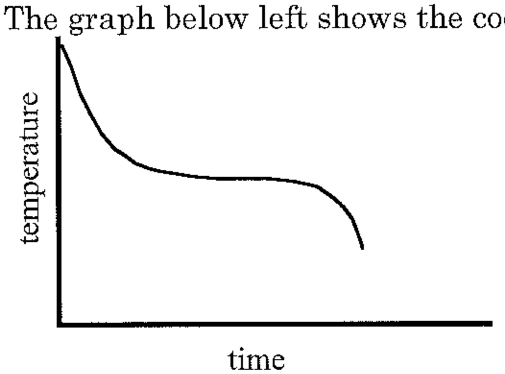
5. One of the ice cubes described in #3 is placed in a glass of room temperature (25 °C) soft drink.
Do separate bar charts for the ice cube and the soft drink.

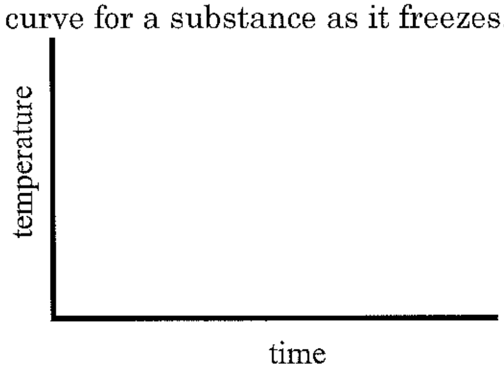




Describe how the arrangement and the motion of the molecules in each system change from the initial to the final state.

6. The graph below left shows the cooling curve for a substance as it freezes.





a. On the graph at right sketch the cooling curve for a larger sample of the same substance.

b. Label which phase (or phases) of the substance is present in each of the three portions of the cooling curve.

c. Describe the arrangement and motion of the molecules during each portion of the graph.

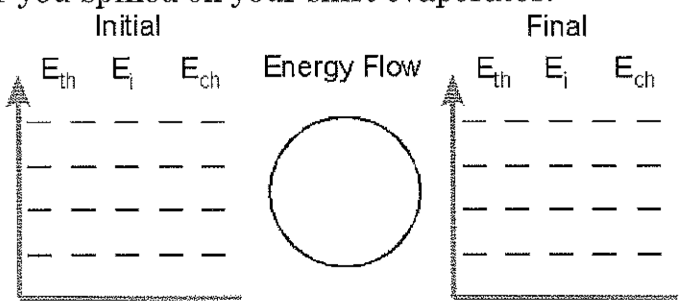
Name________________________________

Date__________ Pd____

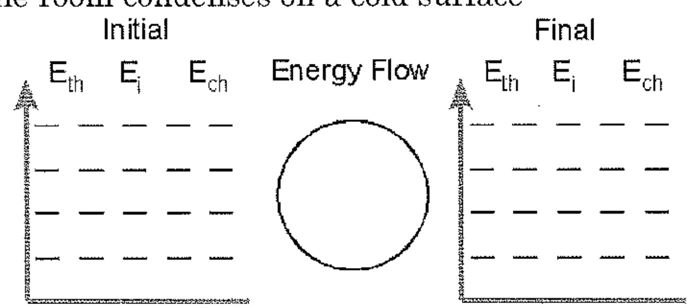
Unit 3 - Worksheet 2

For each of the situations described below, use an energy bar chart to represent the ways that energy is stored in the system and flows into or out of the system. Below each diagram describe how the arrangement and motion of the molecules change from the initial to the final state.

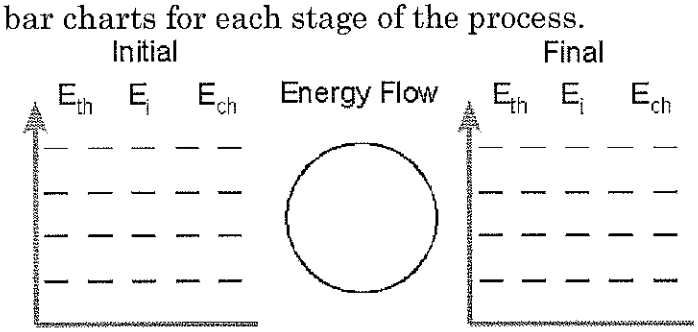
1. Some of the water you spilled on your shirt evaporates.

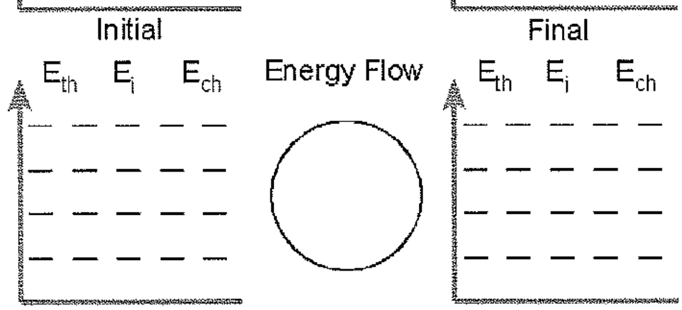


2. Water vapor in the room condenses on a cold surface



3. A pan of water (25˚C) is heated to boiling and some of the water is boiled away. Do separate energy bar charts for each stage of the process.





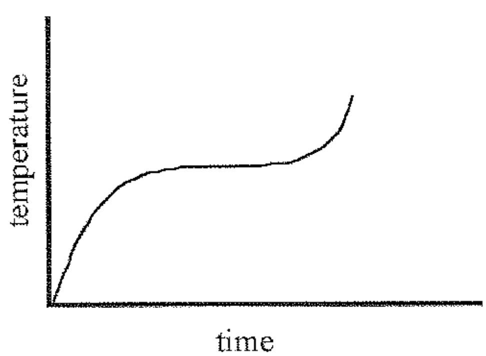
4. During boiling, bubbles appear in the liquid water. In the boxes below represent the arrangement of molecules inside the liquid water and inside a bubble.

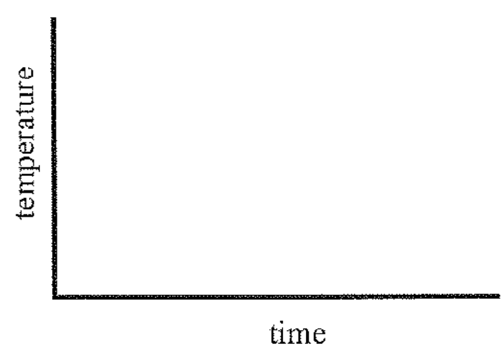
liquid water | bubble

What is inside the bubble? Why do you think so?

5. Suppose the burner under the pan of boiling water is turned to a higher setting. How will this affect the temperature of the water in the pan? Explain.

6. The graph below left represents the heating curve for a liquid heated from room temperature to a temperature above its boiling point.





a. Sketch the heating curve for a larger sample of the same liquid.

b. Label which phase (or phases) of the substance is present in each of the three portions of the heating curve.

c. Describe the arrangement and motion of the molecules during each portion of the graph.